Determination of the Molar Mass of a Volatile Liquid

Introduction

Molar mass is the number of grams of an element or compound in a mole of that substance. Knowing the molar mass of a substance is extremely important. It allows you to converts from grams to moles for a substance, which is integral for stoichiometry. Also knowing the molar mass of a substance is also helpful in determining the identity of that substance. If the unknown substance happens to be volatile, then a variation of the ideal gas law:

$$\text{Molar Mass}=mRT/PV$$

can be used to calculate the molar mass of it.

During the lab, a small sample liquid was evaporated in a test tube. Based on the change of mass of the test tube, the mass of vapor was determined. From this mass, the pressure, volume, and temperature of the gas were substituted into the derivation of ideal gas law above and the molar mass of the substance was calculated.

Methods

A 500 mL beaker was filled with water and brought to a boil on a hot plate to serve as a hot water bath. The mass of a clean, dry test tube and a fitted cork with a small hole in it were massed. Approximately 0.5 mL of the unknown volatile liquid was measured using a pipette and was placed in the test tube. The unknown liquid was colorless and had an odor similar to paint thinner.

Once the hot water bath started to boil, the test tube was stoppered with the cork with a small hole in it and placed into the hot water bath. The tube was completely submerged, with only the cork above the water line. The temperature of the hot water bath, which is the same as the test tube after some time, was recorded. The test tube was allowed to sit in the water bath until the unknown liquid was entirely evaporated. Once all of the unknown liquid had evaporated, the test tube was removed from the water bath and promptly placed in an ice bath, which caused the vapor in the tube to condense. The test tube was dried thoroughly with a paper towel and the mass of the test tube, cork, and condensed gas was then recorded.

The test tube was then rinsed and filled with water, corked (with the same cork used earlier), and massed. This was used to calculate the volume of the test tube. The atmospheric pressure of the room was also recorded.

Results

Table 1 below show the raw data that was collected during the experiment.

Table 1: Raw Data

Mass Empty Test Tube & Cork (g)	20.086
Mass Condensed Gas, Test Tube, & Cork (g)	20.191
Mass Test Tube Filled with H_2O & Cork (g)	47.633
Mass Test Tube Filled with H_2O & Cork (g)	47.633
Mass H_2O in Test Tube (g)	27.547
Temperature (^oC)	96.0
$P_{atm}= P_{test\ tube}$(in Hg)	30.05

Table 2 below shows the data that was calculated during the lab. For calculations, refer to page 72 of the lab notebook.

Table 2: Calculated Data

Mass of Condensed Gas (g)	0.105
Mass H_2O in Test Tube (g)	27.547
Volume of Test Tube (mL)	47.633
Molar Mass of Unknown Liquid (g/mol)	115
Percent Error (%)	98.0

The volume of the test tube was then calculated using the density of water (1.00 g/mL).

Because the cork in the test tube had a small hole in it, gas was allowed to effuse through the cork. Therefore, the pressure inside the test tube is the same as the atmospheric pressure, which was recorded from a weather forecast in inches of mercury. This value had to be converted to torrs (mm Hg), then to atmospheres.

Discussions

The molar mass of the unknown liquid was determined by using the formula, which is derived from the ideal gas law:

$$\text{Molar Mass}=mRT/PV$$

The mass of the gas was found by subtracting the mass of the clean, dry test tube and cork from the mass of the test tube, cork and condensed gas. R is the ideal gas constant, 0.0821 (L atm)/(mol K). T is the temperature of the vapor, in Kelvin, which is assumed to be the same as the temperature of the hot water bath. P is the pressure of the gas in atmospheres, which is assumed to be atmospheric pressure since the container is open to the atmosphere. The pressure of the gas was first measured in inches of mercury. This measurement was multiplied by the number of mm in an inch and divided by the number of mm of Hg that are in one atmosphere. V is the volume of the container, which was determined by dividing the mass of water that filled the test tube by its density. After substituting all of the values into the equation, the molar mass of the unknown liquid was calculated to be 115 g/mol.

The unknown liquid used in this experiment was acetone. Given that acetone (C_3H_6O) has an actual molar mass of 58.1 g/mol, there was a percent error of 98.0%. The experimental molar mass may have been calculated to be too high because the test tube may not have sat long enough in the hot water bath. This would have resulted in less vapor leaving the test tube and having a larger mass of the vapor condensing back into a liquid when cooled. This larger mass of vapor would in turn make the molar mass higher because you would be dividing into a larger number for mass in the equation. In addition, the gas may not have behaved ideally due to its structure. The oxygen atoms cause the acetone molecules to be polar, so they have stronger intermolecular attractions between them than if they were nonpolar. These attractions would potentially cause the molecules to dimerize and result in more molecules of acetone being the test tube during vaporization. This would also increase the mass of liquid and in turn the molar mass. This is more likely the case because the molar mass calculated is about double that of the actual molar mass.

Despite the inaccuracies, the lab outlined a method to determine the molar mass of an unknown liquid. These inaccuracies could have been avoided using a truly nonpolar liquid to avoid issues caused by the dimerization of the polar molecules. Knowing molar mass is useful in determining the identity of a substance. Furthermore. if the empirical formula is also known, then from these the molecular formula could be found.